

Specific Heat

1. A 15.75-g piece of iron absorbs 1086.75 joules of heat energy, and its temperature changes from 25°C to 175°C. Calculate the specific heat capacity of iron.

$$C = \frac{q}{m \Delta T} = \frac{1086.75 \text{ J}}{15.75 \text{ g} (175 - 25)} = \boxed{0.46 \text{ J/g}^\circ\text{C}}$$

2. How many joules of heat are needed to raise the temperature of 10.0 g of aluminum from 22°C to 55°C, if the specific heat of aluminum is 0.90 J/g°C?

$$q = m C \Delta T = 10.0 \text{ g} \left(\frac{0.90 \text{ J}}{\text{g}^\circ\text{C}} \right) (55 - 22) = 297 \rightarrow \boxed{3.0 \times 10^2 \text{ J}}$$

3. To what temperature will a 50.0 g piece of glass raise if it absorbs 5275 joules of heat and its specific heat capacity is 0.50 J/g°C? The initial temperature of the glass is 20.0°C.

$$\Delta T = T_f - T_i = \frac{q}{mC} \quad X - 20 = 5275 \text{ J} \left(\frac{1}{50.0 \text{ g}} \right) \left(\frac{\text{g}^\circ\text{C}}{0.50 \text{ J}} \right)$$

$$X = 211 + 20 = 231 = \boxed{231^\circ\text{C}}$$

4. A certain mass of water was heated with 41,840 Joules, raising its temperature from 22.0 °C to 28.5 °C. Find the mass of the water, in grams. (Cp of H₂O = 4.184 J/g °C)

$$m = \frac{q}{C \Delta T} = \frac{41840 \text{ J}}{\left(\frac{\text{g}^\circ\text{C}}{4.184 \text{ J}} \right) (28.5 - 22)} = 1538 \text{ g} \rightarrow \boxed{1540 \text{ g}}$$

5. What temperature change occurs when 32g of water absorbs 2000 joules of energy? The specific heat capacity of water is 4.18J/g°C.

$$\Delta T = \frac{q}{mC} = \frac{2000 \text{ J}}{32 \text{ g} \left(\frac{\text{g}^\circ\text{C}}{4.184 \text{ J}} \right)} = 14.9^\circ\text{C} \rightarrow \boxed{15^\circ\text{C}}$$

6. What was the initial temperature of a 154g piece of aluminum if the final temperature, after absorbing 3269J, 108°C? The specific heat of aluminum is 0.90 J/g°C.

$$\Delta T = T_f - T_i = \frac{q}{mC} \quad = \quad T_f - \frac{q}{mC} = T_i$$

$$108 - \frac{3269}{154(0.9)} = T_i$$

84.4°C

Calorimetry Problems

1. 240 g of water (initially at 20°C) are mixed with an unknown mass of iron (initially at 500°C). When thermal equilibrium is reached, the system has a temperature of 42°C. Find the mass of the iron. The specific heat of iron is 0.46 J/g°C. (Hint: set $q = mC\Delta T$ for water equal to $q = mC\Delta T$ for iron since the same amount of heat that is lost from the iron will be absorbed by the water)

$$mC\Delta T_w = mC\Delta T_i \quad \frac{m_w C \Delta T_w}{C \Delta T_{Fe}} = m_{Fe}$$

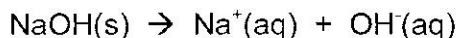
$$240g \frac{4.184J}{g^\circ C} (22^\circ C) \left(\frac{1}{0.46J} \right) \left(\frac{1}{458^\circ C} \right) = 1.0 \times 10^2 g \text{ Fe}$$

2. Compound A is burned in a bomb calorimeter that contains 2.50 liters of water. If the combustion of 0.175 moles of this compound causes the temperature of the water to rise 45.00°C, what is the molar heat of combustion of compound A? The heat capacity of water is 4.184 J/g°C.

$$q = mC\Delta T \quad 2.5 \times 10^3 g \left(\frac{4.184J}{g^\circ C} \right) \left(\frac{45.0^\circ C}{1} \right) = 4.707 \times 10^5 J / 0.175 mol$$

$$1 mol \left(\frac{4.707 \times 10^5 J}{0.175 mol} \right) = \boxed{2.69 \times 10^6 J}$$

3. When a 16.9 g sample of NaOH dissolves in 70.0 g of water in a calorimeter, the temperature rises from 22.4 °C to 86.6°C. Calculate ΔH for the process.



$\Delta H = ?$

(molar heat of solution)

* based on one mole

$$q = mC\Delta T \quad 70.0g \left(\frac{4.184J}{g^\circ C} \right) \left(\frac{64.2^\circ C}{1} \right) = 18802.596$$

$$\frac{18802.596}{16.9g} \left(\frac{40.01g}{1mol} \right) = \boxed{752 KJ}$$

4. Compound B is burned in a bomb calorimeter that contains 1.50 liters of water. When I burned 50.0 grams of compound B in the calorimeter, the temperature rise of the water in the calorimeter was 35.0° C. If the heat of combustion of compound B is 2,150 kJ/mol, what is the molar mass of compound B?

$$q = mC\Delta T \quad 1.5 \times 10^3 g \left(\frac{4.18J}{g^\circ C} \right) \left(\frac{35^\circ C}{1} \right) = 219450 J \rightarrow 219.450 KJ$$

$$MM = \frac{m}{n} \quad \frac{50g}{219.450 KJ} \left(\frac{2150 KJ}{1mol} \right) = \boxed{490.9 g/mol}$$

5. The molar heat of combustion of compound C is 1,250 kJ/mol. If I were to burn 0.115 moles of this compound in a bomb calorimeter with a reservoir that holds 2.50 L of water, what would the expected temperature increase be?

$$q = mC\Delta T \quad 0.115 mol \left(\frac{1250 KJ}{mol} \right) \left(\frac{1000 J}{1 KJ} \right) \left(\frac{1}{2.5 \times 10^3 g} \right) \left(\frac{1}{4.18 J} \right) = \boxed{13.8^\circ C}$$

$$\Delta T = \frac{q}{mC}$$

Heat of Change of State

1. You have a sample of H_2O with a mass of 23.0 g at a temperature of -46.0°C . How many kilojoules of heat energy are necessary to:

a. heat the ice to 0°C ? (hint: no phase change)

$$q_{\text{ice}} = mC\Delta T \quad 23.0\text{g} \left(\frac{2.1\text{J}}{\text{g}^\circ\text{C}} \right) \left(\frac{46^\circ\text{C}}{1} \right) = 2220.8\text{J}$$

b. melt the ice? (hint: phase change)

$$\Delta H_{\text{fus}} = n \times 6.01\text{ kJ/mol} \quad 23.0\text{g} \left(\frac{1\text{mol}}{18\text{g}} \right) \left(\frac{6.01\text{ kJ}}{1\text{mol}} \right) = 7.68\text{ kJ} \text{ or } 7680\text{ J}$$

c. heat the water from 0°C to 100°C ?

$$q_{\text{water}} = mC\Delta T \quad 23.0\text{g} \left(\frac{4.18\text{J}}{\text{g}^\circ\text{C}} \right) \left(\frac{100^\circ\text{C}}{1} \right) = 9610\text{ J}$$

d. boil the water?

$$\Delta H_{\text{vap}} = n \times 40.7\text{ kJ/mol} \quad 23.0\text{g} \left(\frac{1\text{mol}}{18\text{g}} \right) \left(\frac{40.7\text{ kJ}}{1\text{mol}} \right) = 52.0\text{ kJ} \text{ or } 52000\text{ J}$$

e. heat the steam from 100°C to 109°C ?

$$q_{\text{steam}} = mC\Delta T \quad 23.0\text{g} \left(\frac{1.7\text{J}}{\text{g}^\circ\text{C}} \right) \left(\frac{9^\circ\text{C}}{1} \right) = 383.5\text{ J}$$

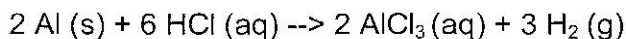
g. Calculate the total energy change from (a-e)

$$\begin{array}{cccccc} 2220.8\text{ J} & + & 7680\text{ J} & + & 9610\text{ J} & + & 52000\text{ J} & + & 383.5\text{ J} \\ \text{a} & & \text{b} & & \text{c} & & \text{d} & & \text{e} \end{array}$$

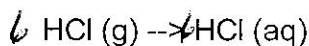
$$\boxed{71900\text{ J}}$$

Hess's Law

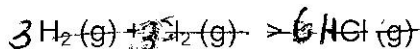
1. Calculate ΔH for the reaction $2 \text{ Al (s)} + 3 \text{ Cl}_2 \text{ (g)} \rightarrow 2 \text{ AlCl}_3 \text{ (s)}$ from the Data.



$$\Delta H = -1049. \text{ kJ}$$

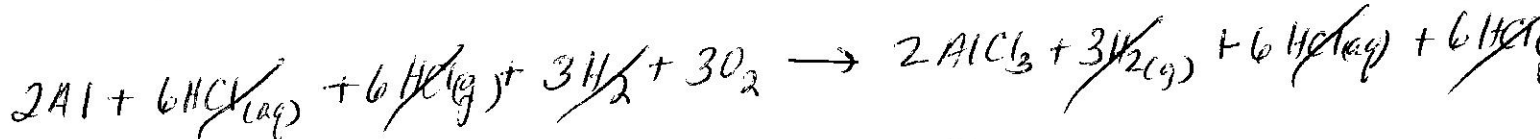


$$\Delta H = -74.8 \text{ kJ} \times 6 = -448.8$$



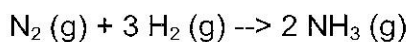
$$\Delta H = -1845. \text{ kJ} \times 3 = -5535$$

$$-7034 \text{ KJ}$$

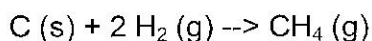


$$\boxed{-7034 \text{ KJ}}$$

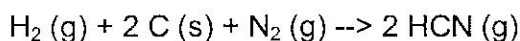
2. Calculate ΔH for the reaction $\text{CH}_4 \text{ (g)} + \text{NH}_3 \text{ (g)} \rightarrow \text{HCN (g)} + 3 \text{ H}_2 \text{ (g)}$, given:



$$\text{flip } \Delta H = -91.8 \text{ kJ} + 91.8$$

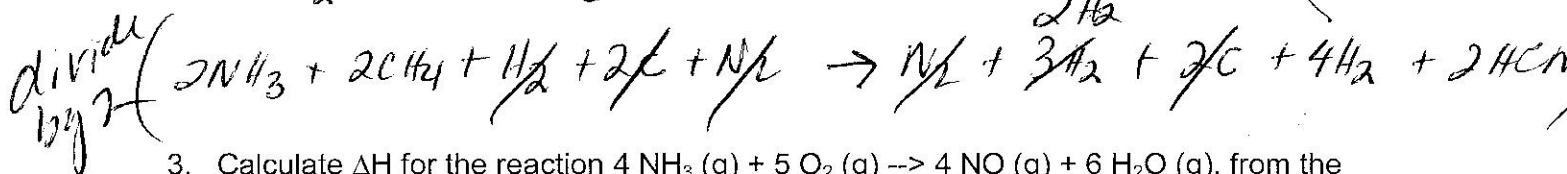
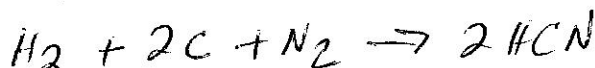
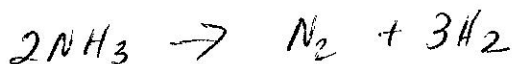


$$\text{flip} \times 2 \Delta H = -74.9 \text{ kJ} + 149.8$$



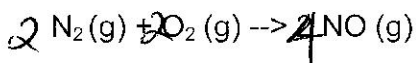
$$\Delta H = +270.3 \text{ kJ} + 270.3$$

$$511.9 \div 2$$

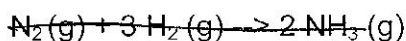


$$\boxed{+256.0 \text{ K}}$$

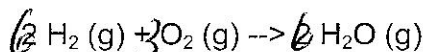
3. Calculate ΔH for the reaction $4 \text{ NH}_3 \text{ (g)} + 5 \text{ O}_2 \text{ (g)} \rightarrow 4 \text{ NO (g)} + 6 \text{ H}_2\text{O (g)}$, from the following Data.



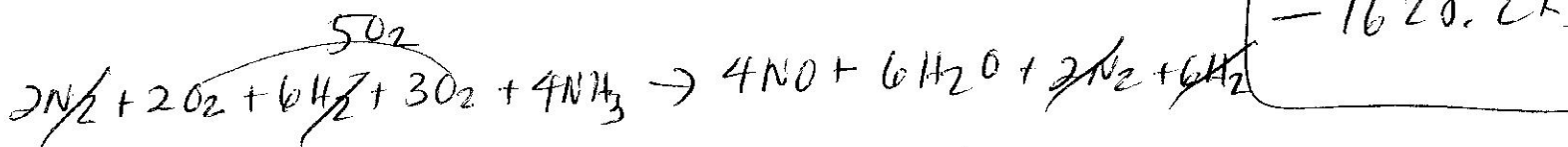
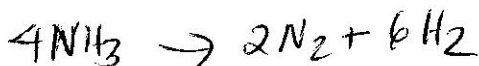
$$\times 2 \Delta H = -180.5 \text{ kJ} = -361$$



$$\text{flip} \times 2 \Delta H = -91.8 \text{ kJ} + 183.6$$



$$\times 3 \Delta H = -483.6 \text{ kJ} - 1450.8$$

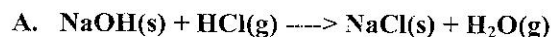


$$\boxed{-1628.2 \text{ K}}$$

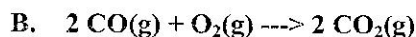
Heat of Formation

$$\Delta H_{rxn} = \sum H_f \text{ products} - \sum H_f \text{ reactants}$$

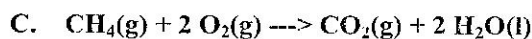
Use a standard enthalpies of formation table to determine the change in enthalpy for each of these reactions.



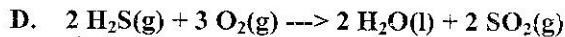
$$(-411 + -241.8) - (-426.7 + -92.3) = -133.8 \text{ KJ}$$



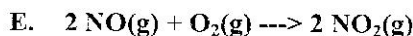
$$(2 \times -393.5) - 2 \text{CO}_2? \text{ Don't have CO}_2 \text{ table got cut off.}$$



$$((2 \times -285.8) + -393.5) - (-74.8) = \boxed{-890.3 \text{ KJ}}$$



$$((2 \times -296.1) + (2 \times -285.8)) - (2 \times -20.1) = \boxed{-1123.6 \text{ K}}$$



$$(2 \times 33.9) - (2 \times 90.4) = \boxed{-113 \text{ KJ}}$$

Compound	ΔH_f (kJ/mol)	Compound	ΔH_f (kJ/mol)
$\text{CH}_4\text{(g)}$	-74.8	HCl(g)	-92.3
$\text{CO}_2\text{(g)}$	-393.5	$\text{H}_2\text{O(g)}$	-241.8
NaCl(s)	-411.0	$\text{SO}_2\text{(g)}$	-296.1
$\text{H}_2\text{O(l)}$	-285.8	$\text{NH}_4\text{Cl(s)}$	-315.4
$\text{H}_2\text{S(g)}$	-20.1	NO(g)	+90.4
$\text{H}_2\text{SO}_4\text{(l)}$	-811.3	$\text{NO}_2\text{(g)}$	+33.9
$\text{MgSO}_4\text{(s)}$	-1278.2	$\text{SnCl}_4\text{(l)}$	-545.2
MnO(s)	-384.9	SnO(s)	-286.2
$\text{MnO}_2\text{(s)}$	-519.7	$\text{SnO}_2\text{(s)}$	-580.7
NaCl(s)	-411.0	$\text{SO}_2\text{(g)}$	-296.1
NaF(s)	-569.0	$\text{SO}_3\text{(g)}$	-395.2
NaOH(s)	-426.7	ZnO(s)	-348.0
$\text{NH}_3\text{(g)}$	-46.2	ZnS(s)	-202.9